CHM129
Acid-Base Equilibrium: Buffers

1. Calculate the pH of a buffer solution that is 0.100 M $HC_2H_3O_2$ and 0.100 M $NaC_2H_3O_2$. $K_a = 1.8 \times 10^{-5}$

$$HC_2H_3O_{2(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + C_2H_3O_2^-{}_{(aq)}$$

	$[HC_2H_3O_2]$	$[H_3O^+]$	$[C_2H_3O_2^-]$
initial	0.100	0.00	0.100
change	$-x$	$+x$	$+x$
equil.	$0.100-x$	x	$0.100+x$

$$\frac{[HA]_{ini}}{K_a} = \frac{0.100}{1.8\times10^{-5}} = 5.6\times10^{3} > 400$$

Assume x is small

$$K_a = \frac{[H_3O^+][C_2H_3O_2^-]}{[HC_2H_3O_2]} = 1.8\times10^{-5}$$

$$\frac{(x)(0.100+x)}{(0.100-x)} = 1.8\times10^{-5}$$

$$\frac{(x)(0.100)}{0.100} = 1.8\times10^{-5}$$

$$x = 1.8\times10^{-5}\,M = [H_3O^+]$$

$$pH = -\log[H_3O^+] = -\log(1.8\times10^{-5})$$

$$\underline{\underline{pH = 4.74}}$$

2. Find the pH, using the Henderson-Hasselbalch equation, of a buffer solution that is 0.100 M $HC_2H_3O_2$ and 0.100 M $NaC_2H_3O_2$. $K_a = 1.8 \times 10^{-5}$

$$pH = pK_a + \log\frac{[base]}{[acid]}$$

$$pH = -\log(1.8\times10^{-5}) + \cancel{\log\frac{(0.100)}{(0.100)}}^{\,0}$$

$$pH = pK_a = \underline{\underline{4.74}}$$

3. A 1.0 L buffer solution contains 0.100 mol $HC_2H_3O_2$ and 0.100 mol $NaC_2H_3O_2$. It has an initial pH = 4.74. The value of K_a for $HC_2H_3O_2$ is 1.8×10^{-5}.

(a) Calculate the new pH after adding 0.010 mol of solid NaOH to the buffer.

(b) For comparison, compute the pH after adding 0.010 mol of solid NaOH to 1.0 L of pure water.

* Ignore any small changes in volume due to the addition of NaOH.

	$HC_2H_3O_{2(aq)}$ +	$NaOH_{(aq)}$ →	$H_2O_{(l)}$ +	$NaC_2H_3O_{2(aq)}$
before addition	0.100 mol	—		0.100 mol
addition	—	0.010 mol		—
after addition	0.090 mol	—		0.110 mol

$$pH = pKa + \log\frac{[base]}{[acid]}$$

$$pH = -\log(1.8\times10^{-5}) + \log\frac{(0.110)}{(0.090)} = \underline{\underline{4.83}}$$

b) $[NaOH] = \dfrac{0.010\text{ mol}}{1.0\text{ L}} = 0.010\text{ M} = [OH^-]$

$pOH = -\log[OH^-] = -\log(0.010) = 2.00$

$pH = 14.00 - pOH = 14.00 - 2.00 = \underline{\underline{12.00}}$